Chapter 10 & 12 Notes- Stoichiometry

**\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_** – A fraction equal to 1 that is used to change one unit into another. (When the numerator = denominator, a fraction equals 1.)

**\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_** – A problem solving method where conversion factors are used to cancel unwanted units.

a) Convert $25 to nickels.

b) Convert 180 days to seconds.

**Common Conversions to Know**

1 base (m, l, g) = 100 \_\_\_\_\_ .

1 base (m, l, g) = 1000 \_\_\_\_ .

1 \_\_\_\_ = 1000 base units (m, l, g)

c) Convert 125 cm to km.

d) Convert 15 m/s to km/hr.

**Measuring Matter**

How do we describe how much of something we have?

By \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_, by \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_, by \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_.

We use words like “dozen” to talk about an amount.

In chemistry, we use the \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_.

**Mole**

**\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_** – SI unit for measuring an amount of a substance.

**A “particle” will either be:**

**An atom, a molecule or a formula unit**

**\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_** = 6.02 x 1023

**\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_** = smallest unit that still has all the characteristics of that substance.

**Representative Particles**

**What is the representative particle of :**

* Element (ex. Cu): \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Exception: The representative particle of the 7 diatomic elements is a molecule. (ex. H2)

* Covalent compound (ex. H2O): \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_
* Ionic Compound (ex. NaCl): \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

4 moles Ca = atoms Ca.

5 x 1018 atoms Cu = moles Cu.

9.2 moles F2 = molecules F2?

9.2 moles F2 = atoms F?

**Molar Mass**

**\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_** – The mass of one mole of an element or compound.

Molar mass of a compound = the sum of the masses of the atoms in the formula.

Use the atomic masses in \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_on the periodic table.

Find the molar mass of each:

Sr

MgBr2

Ba3(PO4)2

**Mole–Gram Conversions**

**1 mole = \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_**

5.3 moles LiOH = \_\_\_\_\_\_\_\_\_\_\_ grams LiOH (Molar mass LiOH : 7 + 16 + 1 = \_\_\_\_\_\_ g/mol)

68 grams F2 = moles F2?

**STP**

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_= Standard Temperature & Pressure

* + Standard Temp 🡪 0oC
  + Standard Press 🡪 1 atm

**Molar Volume of a Gas**

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_: equal volumes of gases at the same temperature and pressure contain equal numbers of particles.

At STP, 1 mole of any gas occupies a volume of \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ L.

\_\_\_\_\_\_\_\_\_\_\_ mole = \_\_\_\_\_\_\_\_\_\_\_\_\_ L (of gas at STP)

5.4 moles He = L He at STP?

**Molar Mass-Density Conversions**

Density = Molar Mass =

A gaseous compound composed of sulfur and oxygen has a density of 3.58 g/L at STP. What is the molar mass of this gas?

What is the density of Krypton gas at STP?

**Multi-step Problem: Example 1**

If you had 5.0 L of CO2 how many grams would that be?

Step 1: L 🡪 moles

Step 2: moles 🡪 grams

x x =

**Multi-step Problem: Example 2**

How many molecules are in 60.0 grams of water?

Step 1: grams 🡪 moles

Step 2: moles 🡪 molecules

**Percent Composition**

**\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_** - % by mass of each element in a compound

Percent = x 100 Percent Composition = x 100

Ex) Find the mass percent composition of Al2(SO4)3

Ex) Find the percent composition of NiSO3

**More Percents**

Ex) Which of the following shows a compound that is 92.3%C and 7.7%H?

a) C2H4 b) C3H6

c) CH4 d) C6H6

**Empirical and Molecular Formulas**

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ – The simplest formula.

Shows the smallest whole number ratio of elements in a compound.

Covalent formulas will not always be empirical.

Example: CH

**\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ Formula** – The actual formula.

For ionic compounds – it will be the simplest ratio.

For molecular compounds – it will NOT always be the simplest ratio.

Example: C6H6

**Empirical Formula**

Calculate the empirical formula of a 2.5 gram compound containing 0.90g Ca and 1.60g Cl.

Step 1: Convert GRAMS to MOLES.

Step 2: DIVIDE the # of moles of each substance by the smallest number to get the simplest mole ratio.

Step 3: If the numbers are whole numbers, use these as the subscripts for the formula. If the numbers are not whole numbers, multiply each by a factor that will make them whole numbers.

Look for these fractions:

What is the empirical formula of a compound that is 66% Ca and 34% P?

Step 1: grams 🡪 moles

Step 2: Divide by the smallest.

**Determining Molecular Formulas**

A compound has an empirical formula of CH2O. Its molecular mass is 180g/mol. What is its molecular formula?

Step 1: Find the mass of the empirical formula.

Step 2: Divide the molecular mass by the mass of the empirical formula to get the “multiplying factor”.

Step 3: Multiply each of the subscripts in the empirical formula by this factor to get the molecular formula.

Find the molecular formula of ethylene glycol (CH3O) if its molar mass is 62 g/mol.

The percent composition of methyl butanoate is 58.8% C, 9.8% H, and 31.4 % O and its molar mass is 204 g/mol.

What is its empirical formula? What is its molecular formula?

Empirical Formula 🡪 C5H10O2

Mass = 5(12) + 10(1) + 2(16) = 102 g/mole

= 2

So molecular formula is 2 x emp. form: 2(C5H10O2) = C10H20O4

**Stoichiometry**

**\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_** – The calculations of quantities of substances involved in chemical reactions

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ in a balanced reaction show the **MOLE TO MOLE RATIO**.

* + Ex. 1N2 + 3H2 🡪 2NH3

There is a mole ratio

Note – Don’t use coefficients when adding up the mass!

**Mole-Mole Conversions**

2A + B 🡪 3C + 7D

Given the number of moles of reactant A (6 moles), you can find:

1)

2)

3)

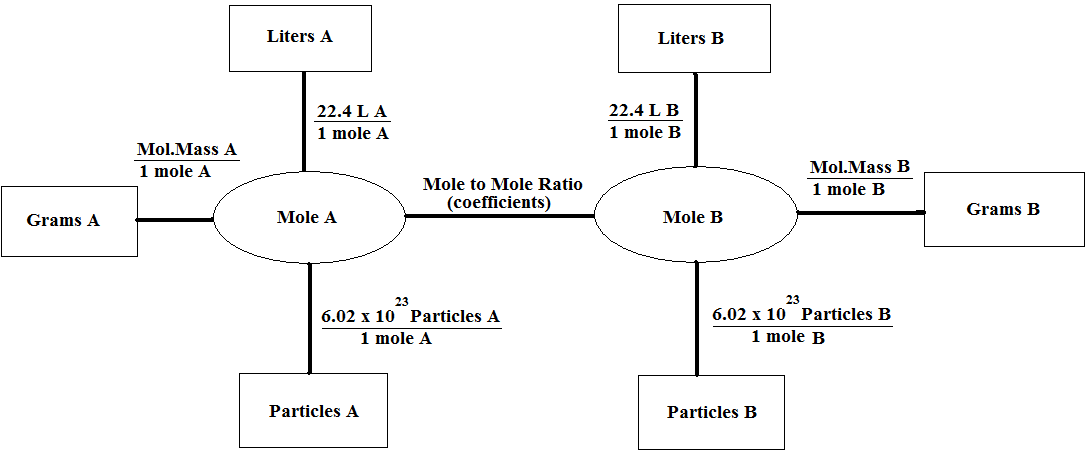
N2 + 3H2 🡪 2NH3

If you have 2 moles of N2, how many moles of NH3 will be produced?

**Reaction Conversions**

\*\*\*\*The only way to convert from one compound to something totally different in the reaction is to use the \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ from the coefficients!!!\*\*\*\*

Note – If you don’t have moles already, your first step is to convert to moles!



**Mass-Mass Conversions**

How many grams of hydrogen gas are required for 3.75g of nitrogen gas to react completely?

(Order: grams N2 🡪 mole N2 🡪 mole H2 🡪 grams H2 )

N2 + 3H2 🡪 2NH3

How many grams of H2 are required to produce 5.0 grams of NH3?

(Grams NH3 🡪 moles NH3 🡪 moles H2 🡪 grams H2 )

**Mass – Volume Conversions**

C2H4 + 3O2 🡪 2H2O + 2CO2

How many liters of CO2 (at STP) would be produced from the combustion of 4.0 grams of C2H4?

C2H4 + 3O2 🡪 2H2O + 2CO2

How many liters of O2 would be needed to produce 6.8 grams of H2O?

**Volume – Volume Conversions**

N2 + 3H2 🡪 2NH3

How many liters of ammonia (NH3) would be produced from 8.0 liters of hydrogen?

N2 + 3H2 🡪 2NH3

How many liters of N2 would be needed to produce 15.0 liters of NH3?